Chapter 6, p. 197 Chemical Proportions in Compounds

Practice Problems p. 201

1. % Ca: $\frac{0.90 \text{ g Ca}}{2.50 \text{ g comound}} \times 100 = 36\%$

% Cl: $\frac{1.60 \text{ g Cl}}{2.50 \text{ g compound}} \times 100 = 64.0\%$

(note that you are working with a compound containing Cl, so you use Cl, not Cl_2. Cl_2 is a molecule of chlorine.)

2. The total mass of the compound is: 7.22 + 2.53 + 5.25 =15.00 g.

% Ni: $\frac{7.22 \text{ g}}{15.00\text{g}} \times 100 = 48.1\%$

% P = $\frac{2.53 \text{ g}}{15.00\text{g}} \times 100 = 16.9\%$

%O = $\frac{5.25}{15.00} \times 100 = 35.0\%$

(again, you are working with a compound containing oxygen, so you are working with O, not O_2 the molecule.)

3. Total mass: 650 mg

mass of H:50.4 mg %H = $\frac{50.4}{650} \times 100 = 7.75\%$

mass of C: 257 mg %C = $\frac{257}{650} \times 100 = 39.5\%$

mass of O = 650 – (50.4 + 257) =342.6 mg %O = $\frac{342.6}{650} \times 100 = 52.7\%$

4. Total mass: 50.0 mg

mass of K:13.3 mg %K = $\frac{13.3}{50.0} \times 100 = 26.6\%$

mass of Cr: 17.7 mg %Cr = $\frac{17.7}{50.0} \times 100 = 35.4\%$

mass of O = 50.0 – (13.3 + 17.7) =19.0 mg %O = $\frac{19.0}{50.0} \times 100 = 38.0\%$

Practice Problems p. 204

5. For each of these % composition questions, you are comparing the number of molar masses (M) of N per molar mass (M) of the compound:

a) N_2O $\frac{2 \text{ M N}}{1 \text{ M } N_2O} \times 100 = \frac{2(14.01) \text{ g}}{44.02 \text{ g}} \times 100 = 63.65\%$

b) $Sr(NO_3)_2$ $\frac{2 \text{ M N}}{1 \text{ M } Sr(NO_3)_2} \times 100 = \frac{2(14.01) \text{ g}}{211.64 \text{ g}} \times 100 = 13.23\%$

c) NH_4NO_3 $\frac{2 \text{ M N}}{1 \text{ M } NH_4NO_3} \times 100 = \frac{2(14.01) \text{ g}}{80.06 \text{ g}} \times 100 = 35.00\%$

d) HNO_3 $\frac{1 \text{ M N}}{1 \text{ M } HNO_3} \times 100 = \frac{1(14.01) \text{ g}}{63.02 \text{ g}} \times 100 = 22.23\%$

6. H_2SO_4 % H: $\frac{2 \text{ M H}}{1 \text{ M } H_2SO_4} \times 100 = \frac{2(1.01) \text{ g}}{98.09 \text{ g}} \times 100 = 2.06\%$

% S: $\frac{1 \text{ M S}}{1 \text{ M } H_2SO_4} \times 100 = \frac{1(32.07) \text{ g}}{98.09 \text{ g}} \times 100 = 32.69\%$

% O: $\frac{4\text{ M O}}{1\text{ M }H_2SO_4} \times 100 = \frac{4(16.00)\text{g}}{98.09\text{ g}} \times 100 = 62.25\%$

7. KNO_3 % O: $\frac{3\text{ M O}}{1\text{ M }KNO_3} \times 100 = \frac{3(16.00)\text{ g}}{101.11\text{ g}} \times 100 = 47.47\%$

8. a) MnO_2 % Mn: $\frac{1\text{ M Mn}}{1\text{ M }MnO_2} \times 100 = \frac{1(54.94)\text{ g}}{86.94\text{ g}} \times 100 = 63.19\ \%$

% O: $\frac{2\text{ M O}}{1\text{ M }MnO_2} \times 100 = \frac{2(16.00)\text{g}}{86.94\text{ g}} \times 100 = 36.81\%$

b) Since you know that MnO2 is 63.19% Mn by mass, then in if you have 250 kg of MnO_2, then you have 0.6319 x 250 kg = 158 kg of Mn.

Section 6.2 p. 207 The Empirical Formula of a Compound

Practice Problems p. 209

In these questions, you go from mass of each element (from % mass) in a compound, to the moles of each element in the compound. Dividing by the smallest number of moles gives the simplest (reduced) mole ratio of each element in the compound, aka the empirical formula, EF.

9.

Element	Mass (g) (from mass percent) in 100 g sample	Molar Mass g/mol (M)	Number of Moles $N = \frac{m}{M}$	Simplest Mole Ratio (÷ smallest n)
H	17.6	1.01	$\frac{1.01}{17.6} = 0.0574$	$\frac{0.0574}{0.0574} = 1$
N	82.4	14.01	$\frac{14.01}{82.4} = 0.170$	$\frac{0.170}{0.0574} = 2.96 = 3$

∴ EF = NH_3

10.

Element	Mass (g) (from mass percent) In 100 g sample	Molar Mass g/mol (M)	Number of Moles $N = \frac{m}{M}$	Simplest Mole Ratio (÷ smallest n)
Li	46.3	6.94	$\frac{46.3}{6.94} = 6.671$	$\frac{6.671}{3.356} = 1.99 = 2$
O	53.7	16.00	$\frac{53.7}{16.00} = 3.356$	$\frac{3.356}{3.356} = 1$

∴ EF = Li_2O

11.

Element	Mass (from mass percent)	Molar Mass (M)	Number of Moles $N = \frac{m}{M}$	Simplest Mole Ratio (÷ smallest n)
B	15.9	10.81	$\frac{15.9}{10.81} = 1.471$	$\frac{1.471}{1.471} = 1$
F	84.1	19.00	$\frac{84.1}{19.00} = 4.426$	$\frac{4.426}{1.471} = 3.009 = 3$

∴ EF = BF_3

12.

Element	Mass (from mass percent)	Molar Mass (M)	Number of Moles $N = \frac{m}{M}$	Simplest Mole Ratio (÷ smallest n)
Cl	52.51	35.45	$\frac{52.51}{35.45} = 1.481$	$\frac{1.481}{1.481} = 1$
S	47.78	32.07	$\frac{47.78}{32.07} = 1.490$	$\frac{1.490}{1.481} = 1.006$

∴ EF = SCl

Practice Problems p. 211

13.

Element	Mass (from mass percent)	Molar Mass (M)	Number of Moles $N = \frac{m}{M}$	Simplest Mole Ratio (÷ smallest n)	Clearing Decimal (multiply by 2)
Cr	68.4	52.00	$\frac{68.4}{52.00} = 1.315$	$\frac{1.315}{1.315} = 1$	1x2 = 2
O	31.6	16.00	$\frac{31.6}{16.00} = 1.975$	$\frac{1.975}{1.315} = 1.50$	1.50 x 2 = 3.00

∴ EF = Cr_2O_3

14.

Element	Mass (from mass percent)	Molar Mass (M)	Number of Moles $N = \frac{m}{M}$	Simplest Mole Ratio (÷ smallest n)	Clearing Decimal (multiply by 2)
P	43.7	30.97	$\frac{43.7}{30.97} = 1.411$	$\frac{1.411}{1.411} = 1$	1x2 = 2
O	56.4	16.00	$\frac{56.4}{16.00} = 3.525$	$\frac{3.525}{1.411} = 2.498$	2.498 x 2 = 5.00

∴ EF = P_2O_5

15.

Element	Mass (from mass percent)	Molar Mass (M)	Number of Moles $N = \frac{m}{M}$	Simplest Mole Ratio (÷ smallest n)	Clearing Decimal (multiply by 2)
Na	17.6	22.99	$\frac{17.6}{22.99} = 0.766$	$\frac{0.766}{0.763} = 1.00$	1.00 x 2 = 2.00
Cr	39.7	52.00	$\frac{39.7}{52.00} = 0.763$	$\frac{0.763}{0.763} = 1$	1 x 2 = 2
O	42.8	16.00	$\frac{42.8}{16.00} = 2.675$	$\frac{2.675}{0.763} = 3.50$	3.5 x 2=7

∴ EF = $NaCr_2O_7$

16.

Element	Mass (from mass percent)	Molar Mass (M)	Number of Moles $N = \frac{m}{M}$	Simplest Mole Ratio (÷ smallest n)	Clearing Decimal (multiply by 3)
C	69.9	12.01	$\frac{69.9}{12.01} = 5.820$	$\frac{5.820}{1.456} = 3.997$	3.997 x 3 = 12.00

H	6.86	1.01	$\frac{6.86}{1.01} = 6.792$	$\frac{6.792}{1.456} = 4.665$	4.665 x 3 = 14.00
O	23.3	16.00	$\frac{23.3}{16.00} = 1.456$	$\frac{1.456}{1.456} = 1$	1 x 3 = 3

∴ EF = $C_{12}H_{14}O_3$

Section Review p. 214

1) a) An **empirical formula** of a compound is also called its simplest formula because the ratios of its constituent atoms are reduced to their smallest numbers.

b) The actual **molecular formula** of a compound will be a whole number multiple (including 1) of the empirical formula. It is therefore possible that an empirical formula is also the molecule's molecular formula.

2.

Element	Mass (%) in 100 g sample	M(g/mol)	n(mol)	mole ratio	mole ratio in whole numbers
C	63.1	12.01	5.25	2.6	8
H	5.31	1.01	5.25	2.6	8
O	31.6	16.00	2.00	1	3

∴ empirical formula of methyl salicylate is $C_8H_8O_3$.

3. a)

Element	Mass (%) in 100 g sample	M (g/mol)	n(mol)	mole ratio	mole ratio in whole numbers
Cl	n/a	n/a	0.315	1	2
O	n/a	n/a	1.1	3.5	7

∴ EF is Cl_2O_7 .

b)

Element	Mass (g) in 100 g sample	M (g/mol)	n(mol)	mole ratio	mole ratio in whole numbers
Si	4.90	28.09	0.17	1	1
Cl	24.8	35.45	0.70	4	4

∴ EF is $SiCl_4$

4.

Element	Mass (g) in 100 g sample (from %)	M(g/mol)	n(mol)	mole ratio	mole ratio in whole numbers
C	40.0	12.01	3.33	1	1

H	6.71	1.01	6.64	2	2
O	53.3	16.00	3.33	1	1

∴ EF is CH_2O for lactic acid.

5.The empirical formula is not the definitive formula of the compound found in my client's possession. It is only the actual formula reduced to its simplest ratio. It does not positively identify the substance in my client's possession. As an example, two very different compounds, one a gas called acetylene with an actual formula of C_2H_2 , and benzene, a liquid with an actual formula of C_6H_6 , both have the same empirical formula of CH. Neither one is positively identified by the empirical formula. The only thing that an empirical formula can prove in a court of law what a formula of a compound is NOT, or what it COULD be.

6.

Element	Mass (g) in 100 g (from %)	M(g/mol)	n(mol)	mole ratio	mole ratio in whole numbers
C	76.54	12.01	6.37	9	9
H	12.13	1.01	12.01	17	17
O	11.33	16.00	0.71	1	1

∴ EF is $C_9H_{17}O$

7.

Element	Mass (g) in 100 g sample (from %)	M(g/mol)	n(mol)	mole ratio	mole ratio in whole numbers
C	74.13	12.01	6.17	5.5	11
H	7.92	1.01	7.84	7	14
O	17.95	16.00	1.12	1	2

∴ EF of phenyl valerate is $C_{11}H_{14}O_2$

8. a)

Element	Mass (g) in 100 g sample (from %)	M(g/mol)	n(mol)	mole ratio	mole ratio in whole numbers
C	64.56	12.01	5.38	10	10
H	5.42	1.01	5.36	10	10
Fe	30.03	55.85	0.54	1	1

$\therefore$ EF of ferrocene is $C_{10}H_{10}Fe$

b) Yes it does. The description says that each molecule of ferrocene contains exactly one atom of Fe. Since the empirical formula has one Fe in it, than the molecular formula of ferrocene is the same as its empirical formula. In addition, if the two rings that 'sandwich' the Fe atom are both the same, than each ring has a formula of C_5H_5 .

Section 6.3, p. 215 <u>The Molecular Formula of a Compound</u>

Practice Problems p. 218

The molecular formula is a whole number multiple (x1, x2, x3… etc) of the empirical formula. The ratio of the molecular formula to the empirical formula is the same as the ratio of the molecular formula mass to the empirical formula mass.

17. Given: E.F. = C_2H_5. $\therefore$ E.F. mass = 2C + 5 H = 2 (12.01) + 5(1.01) = 29.07
Given: M.F. mass = 58 g/mol

$$\frac{\text{M.F.}}{\text{E.F}} = \frac{\text{M.F. mass}}{\text{E.F. mass}}$$

$$\therefore \text{M.F.} = \text{E.F.} \times \frac{\text{M.F. mass}}{\text{E.F. mass}} = C_2H_5 \times \frac{58}{29.07} = C_2H_5 \times 2 = C_4H_{10}$$

18. Given: E.F. = CHO_2. $\therefore$ E.F. mass = C + H + 2 O = (12.01) + (1.01) + 2(16.00)= 45.02
Given: M.F. mass = 90 g/mol

$$\frac{\text{M.F.}}{\text{E.F}} = \frac{\text{M.F. mass}}{\text{E.F. mass}}$$

$$\therefore \text{M.F.} = \text{E.F.} \times \frac{\text{M.F. mass}}{\text{E.F. mass}} = CHO_2 \times \frac{90}{45.02} = CHO_2 \times 2 = C_2H_2O_4$$

19. Given: E.F. = $C_{18}H_{21}NO_3$. $\therefore$ E.F. mass = 18C + 21H + 1N + 3 O
= 18(12.01) + 21(1.01) + 1(14.01) + 3(16.00)= 299.40
Given: M.F. mass = 299 g/mol

$$\frac{\text{M.F.}}{\text{E.F}} = \frac{\text{M.F. mass}}{\text{E.F. mass}}$$

$$\therefore \text{M.F.} = \text{E.F.} \times \frac{\text{M.F. mass}}{\text{E.F. mass}} = CHO_2 \times \frac{299}{299.4} = C_{18}H_{21}NO_3 \times 1 = C_{18}H_{21}NO_3$$

20.

Element	Mass (from mass percent)	Molar Mass (M)	Number of Moles $N = \frac{m}{M}$	Simplest Mole Ratio (÷ smallest n)
C	75.0	12.01	$\frac{75.0}{12.01} = 6.245$	$\frac{6.245}{1.25} = 5$
H	5.05	1.01	$\frac{5.05}{1.01} = 5.00$	$\frac{5.00}{1.25} = 4$
O	20.0	16.00	$\frac{20.0}{16.00} = 1.25$	$\frac{1.25}{1.25} = 1$

$\therefore$ EF = C_5H_4O

$\therefore$ E.F. mass = 5C + 4H + O = 5(12.01) + 4(1.01) + (16.00)= 80.09
Given: M.F. mass = 240.28 g/mol

$$\frac{\text{M.F.}}{\text{E.F}} = \frac{\text{M.F. mass}}{\text{E.F. mass}}$$

$$\therefore \text{M.F.} = \text{E.F.} \times \frac{\text{M.F. mass}}{\text{E.F. mass}} = C_5H_4O \times \frac{240.28}{80.09} = C_5H_4O \times 3 = C_{15}H_{12}O_3$$

Section Review p. 218

1. A mass spectrometer will give you the mass of a molecule of compound in number of protons and neutrons, which is the same as its molar mass. This number, along with the molar mass of the empirical formula, can be used to determine the molecular formula of the compound.

2. $$n\text{ (O atoms)} = \frac{3.61 \text{ X } 10^{24} \text{ atoms O}}{1 \text{ mol tartaric acid}} \times \frac{1 \text{ mol O}}{6.02 \text{ x}10^{23} \text{ atoms O}} = \frac{6 \text{ mol O}}{1 \text{ mol tartaric acid}}$$

Since the empirical formula of $C_2H_3O_3$ has 3 mol of O, the molecular formula of tartaric acid must be $2(C_2H_3O_3) = C_4H_6O_6$

3. A molecular formula identifies exactly how many atoms of each element are in a molecule of a particular substance, providing a unique identification. An empirical formula only give the ratios of the atoms in the molecule in their simplest form. The actual formula is a whole number formula of this. Therefore, an empirical formula simply gives a guide to a 'list' of possible structures without identifying any specific one as the molecule in question.

4. a) Empirical formula is $0.5(C_4H_6O_2) = C_2H_3O$
b) The molar mass of the molecular formula of vinyl acetate is twice that for its empirical formula, since the first formula is twice that of the second.

5. Molecular Formula $C_{6x}H_{5x}O_x$, where x is a whole number multiple
Empirical formula = C_6H_5O
$M(C_6H_5O)$ = 93.11g/mol
Ratio of molar masses = 186 g/mol ÷ 93.11 g/mol = 2
Therefore, molecular formula = $2(C_6H_5O) = C_{12}H_{10}O_2$

Section 6.4, p. 219 <u>Finding Empirical and Molecular Formulas by Experiment</u>

Practice Problems p. 221
These questions are on combustion analysis.

21. Given: 0.539 g sample of a hydrocarbon (i.e. the sample contains only H and C) produces 1.64 g CO_2 and 0.807 g H_2O:

$$\underset{0.539\text{ g}}{C_?H_?} + ?O_2 \rightarrow \underset{1.64\text{ g}}{CO_2} + \underset{0.807\text{ g}}{H_2O}$$

All the grams of C in the 1.64 g of CO_2 came from the $C_?H_?$ sample, and all the grams of H in the 0.807 g H_2O came from the H in the $C_?H_?$ sample.
- To solve this problem you will first determine the grams C and the grams H from the sample.
- Then you will use those masses, to solve the empirical formula of $C_?H_?$.

Step 1: Find the mass of C and H in the sample of $C_?H_?$

The mass of C in 1.64 g CO_2:

$$1.64 \text{ g } CO_2 \text{ x } \frac{1 \text{ M C}}{1 \text{ M } CO_2} = 1.64 \text{ g } CO_2 \text{ x } \frac{1(12.01)C}{1(44.01)\text{g } CO_2} = 0.448 \text{ g C}$$ (mass of carbon before reaction)

The mass of H in 0.807 g H_2O:

$$0.807 \text{ g } H_2O \text{ x } \frac{2 \text{ M H}}{1 \text{ M } H_2O} = 0.807 \text{ g } H_2O \text{ x } \frac{2(1.01)\text{g H}}{1(18.02)\text{ g } H_2O} = 0.0905 \text{ g H}$$

∴ 0.448 g C and 0.0905 g H came from the sample of $C_?H_?$.

Step 2: Use the masses of C and H to solve for the empirical formula of $C_?H_?$

Element	Mass	Molar Mass (M)	Number of Moles $N = \frac{m}{M}$	Simplest Mole Ratio (÷ smallest n)	Clearing the Decimal X 5
C	0.448	12.01	$\frac{0.448}{12.01} = 0.0373$	$\frac{0.0373}{0.0373} = 1$	1 x 5 = 5

H	0.0905	1.01	$\frac{0.0905}{1.01} = 0.0896$	$\frac{0.0896}{0.0373} = 2.4$	2.4 x 5 = 12

$\therefore$ EF = C_5H_{12}

This question also asked for the % composition of the sample.
Given: 0.539 g sample, containing 0.488 g C and 0.0905 g H:

% C: $\frac{0.448 \text{ g C}}{0.539 \text{ g sample}} \times 100 = 83.1\%$

% H: $\frac{0.0905 \text{ g H}}{0.539 \text{ g sample}} \times 100 = 16.8\ \%$

%C + %H = 83.1 + 16.8 = 99.9% which is close enough to 100.

22. This problem is trickier in that you have to solve for the mass of oxygen in the sample as well, then solve for the empirical formula and then also solve for the molecular formula:

$C_?H_?O_? + ?O_2 \rightarrow CO_2 + H_2O$
874 mg 2.23 g 0.652 g
(0.874 g)

Step 1: Find the mass of C and H in the sample of $C_?H_?$
The mass of C in 2.23 g CO_2:

$$2.23 \text{ g } CO_2 \times \frac{1 \text{ M C}}{1 \text{ M } CO_2} = 2.23 \text{ g } CO_2 \times \frac{1(12.01)C}{1(44.01)\text{g } CO_2} = 0.609 \text{ g C}$$

The mass of H in 0.807 g H_2O:

$$0.652 \text{ g } H_2O \times \frac{2 \text{ M H}}{1 \text{ M } H_2O} = 0.652 \text{ g } H_2O \times \frac{2(1.01)\text{g H}}{1(18.02)\text{ g } H_2O} = 0.0731 \text{ g H}$$

$\therefore$ 0.609 g C and 0.0731 g H came from the sample of $C_?H_?O_?$
The oxygen in the CO_2 and the H_2O came from both the $C_?H_?O_?$ and the O_2. The way to find out how much came from just the $C_?H_?O_?$ is:

0.874 g $C_?H_?O_?$ = 0.609 g C + 0.0731 g H + X g O
X g O = 0.874 g $C_?H_?O_?$ – (0.609 g C + 0.0731 g H) = 0.192 g O

$\therefore$ 0.609 g C and 0.0731 g H and 0.192 g O came from the sample of $C_?H_?O_?$

Step 2: Use the masses of C, H and O to solve for the empirical formula of $C_?H_?$

Element	Mass	Molar Mass (M)	Number of Moles $N = \frac{m}{M}$	Simplest Mole Ratio (÷ smallest n)	Clearing the Decimal X 5
C	0.609	12.01	$\frac{0.609}{12.01} = 0.05071$	$\frac{0.05071}{0.0120} = 4.22$	4.22 x 5 = 21
H	0.0731	1.01	$\frac{0.0731}{1.01} = 0.07238$	$\frac{0.07238}{0.0120} = 6.03$	6.03 x 5 = 30
O	0.192	16.00	$\frac{0.192}{16.00} = 0.0120$	$\frac{0.0120}{0.0120} = 1$	1 x 5 = 5

$\therefore$ EF = $C_{21}H_{30}O_5$

Step 3: Finding the molecular formula. Given: molecular mass = 362 g/mol

$\therefore$ E.F. mass = 21C + 30H + 5O = 21(12.01) + 30(1.01) + 5(16.00)= 362.5 g

The empirical formula mass and the molecular formula mass are the same.
$\therefore$ the E.F = the M.F. The molecular formula is $C_{21}H_{30}O_5$

Practice Problems p. 225
These are hydrate problems.

23. Given: the formula, $MgSO_3 \cdot 6H_2O$ and asked for the % by mass water.
This is just a % by mass problem, only you are using a component molecule (H_2O), rather than a component element.

$$\% \text{ mass} = \frac{6 \text{ M } H_2O}{1 \text{ M } MgSO_3 \cdot 6H_2O} \times 100$$

Note, it's the % H_2O in the entire hydrate, so you use the entire formula for the hydrate.

$$= \frac{6(18.02)}{1(212.50)} \times 100 = 50.88\%$$

24. Given: a 3.34 g sample of $SrS_2O_3 \cdot XH_2O$. The mass of the anhydrous salt, SrS_2O_3, is 2.30 g. Find X.
The mass of the water in the sample is: 3.34 g – 2.30 g = 1.04 g
This now becomes an empirical formula problem, but you are working with component molecules (the anhydrous salt and the water), rather than component elements.

Component Molecule	Mass	Molar Mass (M)	Number of Moles $N = \frac{m}{M}$	Simplest Mole Ratio (÷ smallest n)
H_2O	1.04	18.02	$\frac{1.04}{18.02} = 0.05771$	$\frac{0.05771}{0.01151} = 5$
SrS_2O_3	2.30	199.76	$\frac{2.30}{199.76} = 0.01151$	$\frac{0.01151}{0.01151} = 1$

∴ The formula of the hydrate is $SrS_2O_3 \cdot 5H_2O$ (the value of X is 5).

25. Given the formula $Zn(ClO_3)_2 \cdot XH_2O$, with Zn 21.5% by mass
In order to find X, you need the mass of the anhydrous salt, $Zn(ClO_3)_2$, and the mass of the H_2O.

Since the hydrate is 21.5% by mass Zn, then in a 100 g sample of the hydrate, there would be 21.5 g Zn. With the mass of the Zn, you can solve for the mass of the ClO_3.

In the formula for the hydrate, the molar ratio of ClO_3 to Zn, is: $\frac{2 \text{ mol } ClO_3}{1 \text{ mol Zn}}$

mole ratio in a compound has to be the same ratio as molar mass ratio in the compound:

$$\therefore \frac{2 \text{ mol } ClO_3}{1 \text{ mol Zn}} = \frac{2 \text{ M } ClO_3}{1 \text{ M Zn}}$$

Now you can solve for the mass of ClO_3 in 100 g of the hydrate:

$$21.5 \text{ g Zn} \times \frac{2 \text{ M } ClO_3}{1 \text{ M Zn}} = 21.5 \text{ g Zn} \times \frac{2(83.45) \text{ g } ClO_3}{1(65.39) \text{ Zn}} = 54.88 \text{ g } ClO_3.$$

So in a 100 g sample of the hydrate, you have 21.5 g Zn and 54.88 g ClO_3. That means you have: 21.5 + 54.88= 76.38 g of the anhydrous salt $Zn(ClO_3)_2$.

And in the 100 g sample, you have 100- 76.38 = 23.62 g H_2O. Now that you have the mass of the salt and the water, you can solve for X:

Component Molecule	Mass (% in 100 g sample)	Molar Mass (M)	Number of Moles $N = \frac{m}{M}$	Simplest Mole Ratio (÷ smallest n)
H_2O	23.62	18.02	$\frac{23.62}{18.02} = 1.3108$	$\frac{1.3108}{0.3288} = 4$
$Zn(ClO_3)_2$	76.38	232.29	$\frac{76.38}{232.29} = 0.3288$	$\frac{0.3288}{0.3288} = 1$

∴ The formula of the hydrate is $Zn(ClO_3)_2 \cdot 4H_2O$ (the value of X is 4).

Section Review p. 228

2. M ($MgSO_4 \bullet 7H_2O$) = (24.31 + 32.07 +16.00 x 4 + 7(16.00 +2 x 1.01))=246.52 g/mol
M ($MgSO_4$)= 120.38 g/mol

Calculating the mass of Magnesium= $1000 \text{ g Bag} \times \dfrac{24.31 \text{ g Mg}}{246.52 \text{ g } MgSO_4 \bullet 7H_2O} = 98.62 \text{ g Mg}$ in $MgSO_4 \bullet 7H_2O$

Mass of Mg in $MgSO_4 \bullet 7H_2O$ = Mg in $MgSO_4 \bullet 7H_2O$ = 98.62 g

Mass of $MgSO_4$ = $\dfrac{120.38 \text{ g } MgSO_4}{24.31 \text{ g Mg}} \times 98.62 \text{ g Mg} = 488 \text{ g } MgSO_4$

or alternative solution:

Calculating the # moles in 1 kg sample of $MgSO_4 \bullet 7H_2O$: $n = 1000 \text{ g} \times \dfrac{1 \text{ mol}}{246.52 \text{ g}} = 4.056 \text{ mol } MgSO_4 \bullet 7H_2O$

n (moles) of Mg in $MgSO_4$= n of ($MgSO_4 \bullet 7H_2O$)= 4.056 mol

Mass of magnesium: $m(Mg) = 4.056 \text{ mol} \times \dfrac{24.31 \text{ g Mg}}{1 \text{ mol Mg}} = 98.601 \text{ g Mg}$

Mass of magnesium sulphate: $m(MgSO_4) = 98.601 \text{ g Mg} \times \dfrac{120.38 \text{ g } MgSO_4}{24.31 \text{ g Mg}} = 488 \text{ g } MgSO_4$

6. $Zn(NO_3)_2 \bullet x H_2O$

Molar Mass of anhydrous M ($Zn(NO_3)_2$): 65.39 g + 14.01 x 2 + 16.00 x 6 =189.41 g /mol
∴ 1 mol of sample of ($Zn(NO_3)_2$) has 189.41 g of anhydrous zinc nitrate
In 63.67g sample of anhydrous $Zn(NO_3)_2$, there would be 100 g of hydrate
∴ mass of water (H_2O)= 100g – 63.67 g= 36.33 g

Number of moles of anhydrous $Zn(NO_3)$: $n = 63.67 \text{ g} \times \dfrac{1 \text{ mol}}{189.41 \text{ g}} = 0.3361 \text{ mol}$

Number of moles of water H_2O: $n = 36.33 \text{ g} \times \dfrac{1 \text{ mol}}{18.02 \text{ g}} = 2.016 \text{ mol}$

Compound	Number of Moles $N = \dfrac{m}{M}$	Simplest Mole Ratio (÷ smallest n)	Clearning the Decimal
$Zn(NO_3)$	0. 3361 mol	$\dfrac{0.3361 \text{ mol}}{0.3361 \text{ mol}}$	1
H_2O	2.016 mol	$\dfrac{2.016 \text{ mol}}{0.3361 \text{ mol}}$	6

∴ x=6 since the ratio of 1 formula unit of $Zn(NO_3)$ for every 6 molecules of water.

7. 2.524 g sample of carbon, hydrogen, oxygen
a) Mass of carbon (since all the carbon in the sample is converted into CO_2= $3.703 \text{ g } CO_2 \times \dfrac{12.01 \text{ g C}}{44.01 \text{ g } CO_2} = 1.0105 \text{ g C}$ in sample

Mass of hydrogen in sample (since all the hydrogen in sample is converted into H_2O= 1.514 g H_2O x $\frac{2.02 \text{ g H}}{18.02 \text{ g } H_2O}$ = 0. 1697 g H_2 in sample

Mass of oxygen in sample= 2.524 g – 1.0105 g – 0.1697 g = 1.3438 g O

Element	Mass (g)	Molar Mass (M) g/mol	Number of Moles $N = \frac{m}{M}$	Simplest Mole Ratio (÷ smallest n)
C	1.0105	12.01	$\frac{1.0105}{12.01} = 0.08414$	$\frac{0.08414}{0.08399} = 1$
H	0.1697	1.01	$\frac{0.1697}{1.01} = 0.1680$	$\frac{0.1709}{0.08399} = 2$
O	1.3438	16.00	$\frac{1.3438}{16.00} = 0.08399$	$\frac{0.08399}{0.08399} = 1$

∴ Empirical formula of the compound containing carbon, hydrogen and oxygen is CH_2O

b) Molecular formula 6 x CH_2O = $C_6H_{12}O_6$

Chapter Review p. 229

1. This simplifies calculations, as the molar masses of the compound, and that of the elements can be used to calculate percentage composition data.

2. It is exactly the same as water found in nature, as all water has the same formula, H_2O, regardless of how it was created in the first place.

3. a) Masses of CO_2 and H_2O generated during the combustion, as well as the mass of the original hydrocarbon burned.
   b) Their results would be the same, and the C-H combustion analyzer provides data that can be used to calculate an empirical formula, not a molecular formula. Additional information obtained through other means is needed to calculate a molecular formula.

4. No, there is not enough information to get an empirical formula in the first place, as the mole ratios of the elements must be known.

5. $M(Na_2B_4O_7)\bullet 10H_2O$ = 381.42 g/mol
   $n((Na_2B_4O_7)\bullet 10H_2O)$ = 5.00 g ÷ 381.42 g/mol = 0.0131 mol borax $(Na_2B_4O_7)\bullet 10H_2O$
   therefore $n(H_2O)$ = 10 X 0.0131 mol = 0.131 mol
   $m(H_2O)$ = 0.131 mol X 18.02 g/mol = 2.36 g
   therefore mass of borax remained after water was removed:
   $m(Na_2B_4O_7)$ = 5.00 g - 2.36 g = 2.64 g

6. a) $M(CCl_2F_2)$ = 120.91 g/mol
   %cm (C) = 12.01 g/mol ÷ 120.91 g/mol X 100% = 9.93 %
   %cm (Cl) = 2 X 35.45 g/mol ÷ 120.91 g/mol X 100% = 58.6%
   %cm (F) = 2 X 19.00 g/mol ÷ 120.91 g/mol X 100% = 31.4%

   b) $M[Pb_3(OH)_2(CO_3)_2]$ = 776.15 g/mol
   %cm (Pb) = 3 X 207.37 g/mol ÷ 776.15 g/mol X 100% = 80.2%
   %cm (O) = 8 X 16.00 g/mol ÷ 776.15 g/mol X 100% = 16.5%
   %cm (H) = 2 X 1.01 g/mol ÷ 776.15 g/mol X 100% = 0.260%
   %cm (C) = 2 X 12.01 g/mol ÷ 776.15 g/mol X 100% = 3.09%

7. a) $M(MgCl_2C_2H_2O)$ = 131.25 g/mol
   $n(MgCl_2C_2H_2O)$ = 25.00g ÷ 131.25 g/mol = 0.1905 mol
   $n(H_2O)$ = 2 X 0.1905 mol = 0.3810 mol
   $m(H_2O)$ = 0.3810 mol X 18.02 g/mol = 6.866 g

b) $M(KMnO_4)$ = 158.04 g/mol
%cm (Mn) = 54.94 g/mol ÷ 158.04 g/mol X 100% = 34.76%
m(Mn) = 34.76% X 5.00 g = 1.74 g

8. a) $M(AgNO_3)$ = 169.88 g/mol
%cm (Ag) = 107.87 g/mol ÷ 169.88 g/mol X 100% = 63.50%
b) m(Ag) = 63.50% 2.00 X 102 kg = 127 kg

9. $M(BaSO_4)$ = 233.40 g/mol
%cm (Ba) = 137.33 g/mol ÷ 233.40 g/mol X 100% = 58.84%
m(Ba) = 58.84% X 45.8 g = 26.9 g

10. $M[Bi(NO_3)_2]$ = 333.00 g/mol
%cm (Bi) = 208.98 g/mol ÷ 333.00 g/mol X 100% = 62.76%
m(Bi) = 62.76% X 268 g = 168 g

11. M(Empirical) = 30.03 g/mol
M(Molecular) ÷ M(Empirical) = 121 g/mol ÷ 30.03 g/mol = 4
Therefore, molecular formula is $4(CH_2O) = C_4H_8O_4$

12. $M(C_6H_2OCl_2)$ = 160.98 g/mol
M(Molecular) ÷ M(Empirical) = 322 g/mol ÷ 160.98 g/mol = 2
Therefore, molecular formula is $2(C_6H_2OCl_2) = C_{12}H_4O_2Cl_4$

13. a) A formula cannot have a fractional subscript, as this implies fractional atoms in a formula.
b) empirical formula = $3(CH_{2.67}) = C_3H_8$

14.

Element	Mass (%)	M(g/mol)	n(mol)	mole ratio	mole ratio in whole numbers
C	80.2	12.01	6.67	10.5	21
H	9.62	1.01	9.54	15	30
O	10.18	16.00	0.636	1	2

Therefore empirical formula is $C_{21}H_{30}O_2$

15.

Element	Mass (%)	M(g/mol)	n(mol)	mole ratio	mole ratio in whole numbers
Na	17.6	22.99	0.766	1	2
Cr	39.7	52.00	0.763	1	2
O	42.8	16.00	2.67	3.5	7

Therefore empirical formula is $Na_2Cr_2O_7$

16.

Element	Mass (%)	M(g/mol)	n(mol)	mole ratio	mole ratio in whole numbers
Hg	67.6	200.59	0.337	1	1
S	10.8	32.07	0.337	1	1
O	21.6	16.00	2.00	4	4

Therefore empirical formula is $HgSO_4$

17. a)

Element	Mass (%)	M(g/mol)	n(mol)	mole ratio	mole ratio in whole numbers
Ca	38.8	40.08	0.968	1.5	3
P	20.0	30.97	0.646	1	2
O	41.2	16.00	2.575	4	8

b) Since each formula unit as 2 P, and so does the empirical formula, the molecular formula is most likely the same. It is more recognizable as $Ca_3(PO_4)_2$.

18. a)

Element	Mass (%)	M(g/mol)	n(mol)	mole ratio	mole ratio in whole numbers
C	71.0	12.01	5.91	18	18
H	8.60	1.01	8.51	26	26
O	15.8	16.00	0.988	3	3
N	4.60	14.01	0.328	1	1

Therefore, the empirical formula of Capsaicin is $C_{18}H_{26}O_3N$

b) Since both the empirical and molecular formulas contain one N, the molecular formula is the same as the empirical formula $C_{18}H_{26}O_3N$

19. Create a table as for empirical formula and work backwards to determine the molar mass of X, therefore identifying it.

Element	Mass (%)	M(g/mol)	n(mol)	mole ratio	mole ratio in whole numbers
X	56.00	??????	1.1	2	2
O	44.00	16.00	2.75	5	5

Therefore, M = 56.00g ÷ 1.1 mol = 50.91 g/mol
V, with a M=50.94 g/mol, is the closest. A formula of V_2O_5 is a valid formula!

20. m(H from HCl) = 4.730g X 1.01g/mol ÷ 36.46 g/mol = 0.131 g
m(C from CCl_4) = 9.977g X 12.01 g/mol ÷ 153.81 g/mol = 0.7790 g
m(O) = 1.254 g - 0.131 g - 0.7790 g = 0.344 g

Element	Mass (g)	M(g/mol)	n(mol)	mole ratio	mole ratio in whole numbers
C	0.7790	12.01	0.06486	3	3
H	0.131	1.01	0.130	6	6
O	0.344	16.00	0.0215	1	1

Therefore, the empirical formula is C3H6O

21. $m(FeSO_4)$ = 1.52 g
   $n(FeSO_4)$ = 1.52 g ÷ 151.92 g/mol = 0.0100 mol
   $m(H_2O)$ = 2.78 g - 1.52 g = 1.26 g
   $n(H_2O)$ = 1.26 g ÷ 18.02 g/mol = 0.0699
   mole ratio = 7:1
   Therefore, there are 7 water molecules for every formula unit of iron(II) sulfate.

22. a) m(C) from CO2 = 0.6871 g X 12.01 g/mol ÷ 44.01 g/mol = 0.1875 g
   m(H) from H2O = 0.1874 g X 2.02 g/mol ÷ 18.02 g/mol = 0.02101 g
   m(O) = 0.5000 g - 0.1875 g - 0.02101 g = 0.2915 g
   %cm (C) = 0.1875 g / 0.5000 g X 100% = 37.50%
   %cm (H) = 0.02101 g / 0.5000 g X 100% = 4.202%
   %cm (O) = 0.2915 g / 0.5000 g X 100% = 58.30%

b)

Element	Mass (g)	M(g/mol)	n(mol)	mole ratio	mole ratio in whole numbers
C	0.1875	12.01	0.01561	1	6
H	0.02101	1.01	0.02081	1.33	8
O	0.2915	16.00	0.01822	1.167	7

Therefore, empirical formula of citric acid is $C_6H_8O_7$

c) M(empirical) = 192.14 g/mol
   M(molecular) = 192 g/mol
   Therefore, molecular formula of citric acid is $C_6H_8O_7$.

23. M(methanol) = 32.05 g/mol
   %cm (C) = 12.01 g/mol ÷ 32.05 g/mol X 100% = 37.47%
   %cm (H) = 4.04 g/mol ÷ 32.05 g/mol X 100% = 12.61%
   %cm (O) = 16.00 g/mol ÷ 32.05 g/mol X 100% = 49.92%
   m(C) in methanol = 1.00g X 37.47% = 0.3747 g
   m(H) in methanol = 1.00g X 12.61 % = 0.1261 g

   therefore,
   $m(CO_2)$ = 0.3747g X 44.01 g/mol ÷ 12.01 g/mol = 1.373 g
   $m(H_2O)$ = 0.1261 g X 18.02 g/mol ÷ 2.02 g/mol = 1.125 g